Unit 5 – Counting Particles - Objectives

Review Concepts a. Types of substances b. Chemical formulas of substances (U 4)	
1. State evidence for Avogadro's Hypothesis. Use Avogadro's Hypothesis and experimental data to determine the relative mass of molecules.	
2. Use experimental data to determine the relative mass of two objects.	
. Use experimental data to determine the number of items in a sample without actually counting them.	
4. Given the chemical formula of a substance, determine the molar mass.	
5. Given the mass of a substance, determine a. the number of moles of the sample b. the number of atoms or molecules in the sample	

6. Given the number of moles of a substance, find a. the mass of the sample b. the number of atoms or molecules in the sample	
7. Given the formula of a compound, determine its % composition.	
8. Given data about the % composition of a sample, determine the empirical formula of the compound.	
9. Given the empirical formula and information about the molar mass of the compound, determine the molecular formula.	
Vocabulary • Avagadro's Hypothesis • relative mass • mole • percent composition • empirical formula • Avagadro's number	

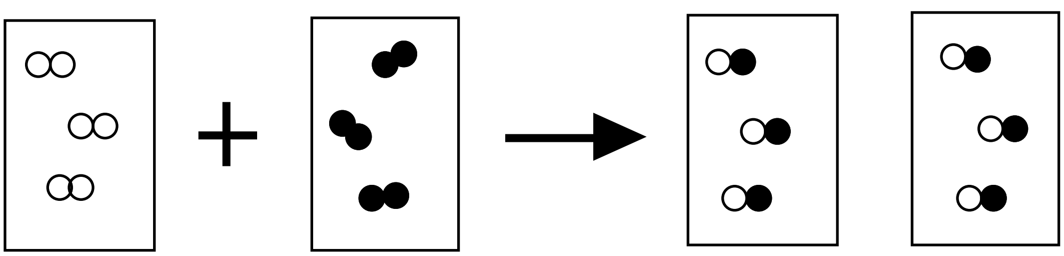
Unit 5 - Avogadro's Hypothesis

Avogadro built on the work of Gay-Lussac who first noted that gases reacted in simple integer volume ratios (when measured at the same temperature and pressure). Gay-Lussac concluded that this result could be explained if the volumes contained the *same number of particles*. Chemists found this hypothesis difficult to accept because they reasoned, for example, that if the particles of the gases were combining, then one volume of gas A reacting with one volume of gas B should produce one volume of product.

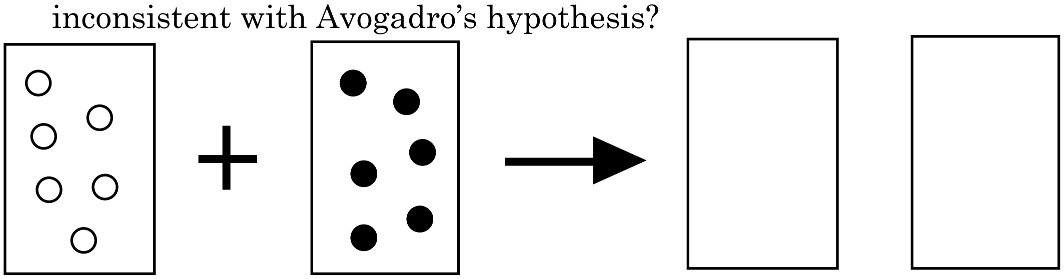
However, it was frequently found that one volume of gas A reacted with one volume of gas B to produce *two volumes* of product. Chemists, including Gay-Lussac, were unable to account for this behavior of gases.

Avogadro reasoned that these results could be explained if Gay-Lussac's particles were actually combinations of smaller particles (which we now call "atoms"). He introduced the use of the term "molecule" to refer to these combinations of smaller particles. "Molecule" comes from "mole" – lumps of matter and "cula" – little. A molecule is a little lump of matter. His hypothesis was consistent with the view that gas pressure is caused by the collisions of molecules with the side of the container. If the pressure is the same, then it seems reasonable that the number of molecules in the container is the same.

In addition to the ammonia and hydrogen chloride example combining in a one to one ratio, he found that a number of gases combined in such a way that could only be accounted for if the individual molecules were diatomic. The diagram below shows how one volume of hydrogen reacts with one volume of chlorine to produce two volumes of hydrogen chloride. Making the assumption that the gases are diatomic neatly accounts for the experimental evidence.



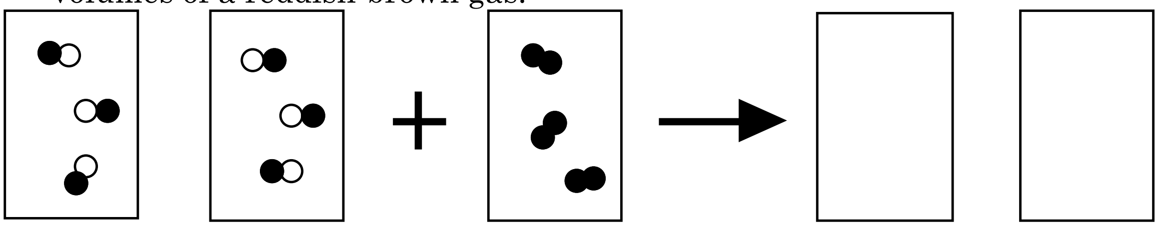
Suppose that these gases were not diatomic. In the "product containers" combine the atoms of hydrogen and chlorine to make hydrogen chloride. Why is your representation inconsistent with Avogadro's hypothesis?



Of course, Avogadro did not base his hypotheses about number of molecules in a volume and the diatomic nature of molecules on just a couple examples. He cited the results of many experiments. Still, the scientific community was not able to accept the validity of his hypothesis until many years after his death.

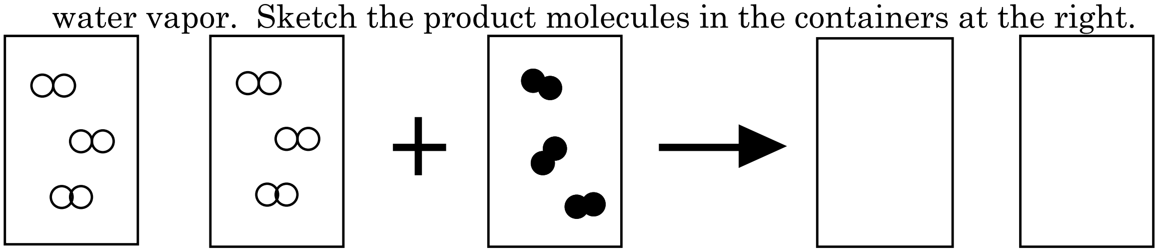
Using his hypothesis, chemists were able to deduce the formulas of gaseous substances based on combining volumes. The use of this hypothesis was also instrumental in determining the molar mass of elements.

Consider the case of two colorless gases, nitric oxide and oxygen gas, reacting to form two volumes of a reddish-brown gas.



Sketch the product molecules in the two containers after the reaction arrow. Explain how were you able to deduce the formula of these molecules.

Two volumes of hydrogen gas combine with one volume of oxygen gas to form two volumes of water vapor. Sketch the product molecules in the containers at the right.



Dalton insisted until his death that the formula of water was OH. Why do we believe that the correct formula for water is H_2O?

Why must oxygen gas be diatomic?

Name ______________________

Date______________ Pd ___

Activity: Relative Mass

Purpose

The purpose is to determine the relative mass of different kinds of hardware and to learn to count by massing.

Data

Hardware	Mass (g)
Empty vial	
Vial + Washers	
Vial + Hex Nuts	
Vial + Bolts	

Calculations

Set up the calculations on the back of this paper. Be sure to label your quantities!

1. Each vial contains the same number of pieces of hardware. Calculate the *relative* mass of each kind of hardware. Divide each mass by the mass of the smallest. (The smallest will be 1.00)

2. A box of hardware contains 100 pieces. Assuming there are 25 pieces in each vial, calculate the mass of a box of each kind of hardware. Express these values in units of g/box.

3. If you had 1.00 kg of each kind of hardware, how many boxes of each would you have?

4. You learned that a barrel of the 1" bolts had a mass of 65.2 kg. The mass of the barrel was 9.6 kg. How many boxes of bolts are in the barrel?

5. Someone at the Home Depot tells you that a 2" bolt is 6.75 times as heavy as a washer. What would be the mass of a box of such bolts?

Conclusion

Suppose that you were given the job of shipping 25,000 hex nuts to a customer. How many boxes of hex nuts would this be? All you have is a hanging scale and a barrel of hex nuts. Describe how you could determine the proper number of pieces without physically counting them out.

Counting in Chemistry

Using mass to count
&
Grouping particles into manageable chunks

Say that a jellybean has a mass of 1.333 g

You can write that as 1.333 g/jellybean.

What does the "/" mean?

If you had a dozen jellybeans, how many jellybeans do you actually have? Of course you have 12!

You can write that as 12 jellybeans/Dozen.

What is the mass of a 2.00 dozen jellybeans? (group→ mass conversion)

If you have 80 g of jellybeans, how many dozen jellybeans do you have? (mass → group conversion)

If you have 160 g of jellybeans, how many jellybeans do you have? (mass → particle conversion)

If you have 78 jellybeans, how many dozen jellybeans do you have?
(particle→ group conversion)

How large a group is a dozen atoms or molecules?

An average atom is about 10^{-10} m wide. In other words, if you put 10 BILLION atoms in a line, they would stretch 1 meter.

How long would a line of a dozen atoms be?

Common Groups of Objects

Dozen = 12 particles

Case = 24 particles

Gross = 144 particles

Ream = 500 particles

Mole = 6.02×10^{23} particles

Googol = 1×10^{100} particles

Why do we needs groups like a "mole" or a "googol"?

Check out Page 6 of your packet – "The Mole"

The Mole and the Periodic Table

The masses listed for each element are the "MOLAR MASSES" for that element.

A "molar mass" is the mass of 1.00 moles of that element.

Ex. Hydrogen has a molar mass of 1.0079 g/mole

Oxygen has a molar mass of 15.999 g/mole

Water has a molar mass of 18.0 g/mole

Chemistry – Unit 5 Notes

1. State Avogadro's Hypothesis.

   What evidence supports this hypothesis?

   Why do chemists find it so useful?

2. What is relative mass? If the mass of A is 5.0 g and the mass of B is 80 g, what are their relative masses?

3. What is molar mass? Why does it allow the chemist to do?

4. How do you find the molar mass of a compound? Show an example.

5. Show an example of how to convert mass (in g) to moles.

6. How many particles in a mole?

7. Show a couple of examples in which you calculate the number of molecules or atoms in a sample of a compound.

8. What is % composition? Show an example in which you calculate % composition.

9. What is empirical formula? Show an example in which you calculate the empirical formula of a compound.

10. What is molecular formula? Show an example in which you calculate the molecular formula of a compound.

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Chemistry Unit 5 – The Mole

To help you better visualize the enormous size of Avogadro's number, 6.02×10^{23}, consider the following analogies:

1. If we had a mole of rice grains, all the land area of the earth would be covered with rice to a depth of about 75 meters!

2. One mole of rice grains is more grain than the number of **all** grain grown since the beginning of time.

3. One mole of marshmallows (standard 1 in^3 size) would cover the United States to a depth of 650 miles.

4. If the Mount St. Helens eruption had released a mole of particles the size of sand grains, the entire state of Washington would have been buried to a depth equal to the height of a 10-story building.

5. A mole of basketballs would just about fit perfectly into a ball bag the size of the earth.

Your turn

Show your solutions to the following questions on the back of this sheet. Multiply by factors and show the cancellation of units. Keep 2 sf's in your answers.

6. Assuming that each human being has 60 trillion body cells (6×10^{13}) and that the earth's population is 6 billion (6×10^{9}), calculate the total number of living human body cells on this planet. Is this number smaller or larger than a mole? Divide the larger value by the smaller to determine the relative size of the two values.

7. One of the fastest supercomputers can perform about 12 teraflops (1 teraflop is 10^{12} calculations per second). Determine how many seconds it would take this computer to count a mole of things. Convert this figure into years.

8. If you started counting when you first learned how to count and then counted by ones, eight hours a day, 5 days a week for 50 weeks a year, you would be judged a 'good counter' if you could reach 4 billion by the time you retired at age 65. If every human on earth (about 6×10^{9}) were to count this way until retirement, what fraction of a mole would they count?

Chemistry Unit 5 - Empirical Formula Lab

Introduction

In this experiment, a measured amount of zinc will be allowed to react with hydrochloric acid, HCl. One of the reaction products will be zinc chloride. You will obtain data that will enable you to determine the empirical formula of zinc chloride, Zn_xCl_y. Empirical means “based on experimental evidence”.

Procedure - Day 1

1. Find the mass of a clean, dry, labeled beaker (to the nearest 0.01g)

2. Add the number of zinc pieces as instructed by your teacher. Find the mass of the beaker and zinc to the nearest 0.01g.

3. Add 50 mL of 3M HCl. Record your observations.

4. Place your labeled beaker on one of the hot plates in the fume hood.

Procedure - Day 2

5. Set up a bunsen burner, ring stand, ring and wire screen so you can heat the beaker.

6. Zinc chloride absorbs water readily from the air. In order to remove the water, heat the beaker and contents for a minute or two. As long as the contents appear to bubble, water is evaporating. However, when the contents begin to smoke, stop heating *immediately*. Remove the beaker (use tongs, it's hot) and allow it to cool on the metal base of the ring stand. Note how the zinc chloride solidifies from the molten state.

7. When the beaker is cool enough to handle, find the mass of the beaker and zinc chloride. (1)

8. Repeat steps 6 and 7. (2)
   a. If this second mass is more than 0.02g lighter than the previous mass, repeat steps 6 and 7 once more. (3, if needed)
   b. If the mass is unchanged, wash out the contents of the beaker as instructed by your teacher.

Data

Mass of labeled beaker ____________ g

Mass of beaker + zinc ____________ g

Mass of beaker + zinc chloride (1) ____________ g

Mass of beaker + zinc chloride (2) ____________ g

Mass of beaker + zinc chloride (3, if needed) ____________ g

Observations

Calculations

1. Determine the mass of zinc reacted.

2. Determine the mass of zinc chloride (guess which one you should use).

3. Determine the mass of chlorine in the zinc chloride.

4. Determine the number of moles of zinc, then the number of moles of chlorine.

5. Determine the ratio: $\dfrac{moles\, Cl}{moles\, Zn}$

Conclusion

1. Since you believe that atoms combine in simple, whole-number ratios, what do you think is the likely ratio: $\dfrac{atoms\, Cl}{atoms\, Zn}$?

2. How does your value compare to the accepted value?

3. What is the empirical formula of zinc chloride?

4. Suppose that you had not driven off all the water from the zinc chloride. How would this error have affected the ratio in calculation 5? Show evidence for your prediction by repeating calculations 2 – 5 using the next to the last value for the mass of the beaker and zinc chloride.

Name ____________________

STOICHIOMETRY PROBLEMS

Moles of Elements—One-Step Problems (continued)

Exercises

Begin each problem by sketching a diagram that outlines the steps in the solution to the problem.

Convert to moles.

1. 12.04×10^{23} atoms He **1.** ____________

2. 3.01×10^{23} atoms Cu **2.** ____________

3. 3.612×10^{23} atoms Fe **3.** ____________

4. 100 atoms Ar **4.** ____________

5. 1 atom S **5.** ____________

6. 24 grams C **6.** ____________

7. 59.3 grams Sn **7.** ____________

8. 98.9 grams Na **8.** ____________

9. 5000 grams K **9.** ____________

10. 0.005 00 gram Ne **10.** ____________

Convert to mass in grams.

11. 10.0 moles Na **11.** ____________

12. 2.20 moles Sn **12.** ____________

13. 5.00 moles Ag **13.** ____________

14. 0.000 300 mole Au **14.** ____________

15. 1.00×10^{7} moles B **15.** ____________

Convert to number of atoms.

16. 3.00 moles Ar **16.** ____________

17. 8.50 moles Fe **17.** ____________

18. 25.0 moles Ar **18.** ____________

19. 0.001 00 mole Na **19.** ____________

20. 1.0×10^{-5} mole Al **20.** ____________

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Moles of Compounds—One-Step Problems (continued)

Exercises

Begin each problem by sketching a diagram that outlines the steps in the solution to the problem.

Convert to moles.

1. 6.02×10^{23} molecules CO_2 — 1. ________
2. 1.806×10^{23} molecules Cl_2 — 2. ________
3. 1.51×10^{23} molecules H_2O — 3. ________
4. 1000 molecules P_4O_{10} — 4. ________
5. 1 molecule NH_3 — 5. ________
6. 34 grams NH_3 — 6. ________
7. 50.0 grams $CaCO_3$ — 7. ________
8. 360 grams H_2O — 8. ________
9. 9.00 grams H_2SO_4 — 9. ________
10. 1.00 gram NaCl — 10. ________

Convert to mass in grams.

11. 5.0 moles NH_3 — 11. ________
12. 4.50 moles NaCl — 12. ________
13. 0.30 mole HCl — 13. ________
14. 0.002 00 mole Na_2SO_4 — 14. ________
15. 1.50×10^{-4} mole AgCl — 15. ________

Convert to number of molecules.

16. 2.0 moles CO_2 — 16. ________
17. 1.8 moles PCl_3 — 17. ________
18. 35.0 moles NH_3 — 18. ________
19. 0.0500 mole SO_2 — 19. ________
20. 1.00×10^{-3} mole CO — 20. ________

10B

Name ______________________________

STOICHIOMETRY PROBLEMS

Moles of Elements—Two-Step Problems (continued)

Exercises

Begin each problem by sketching a diagram that outlines the steps in the solution to the problem.

Convert to mass in grams.

1. 6.02×10^{23} atoms Ca **1.** __________

2. 1.204×10^{23} atoms Bi **2.** __________

3. 3.01×10^{23} atoms Ni **3.** __________

4. 1000 atoms Al **4.** __________

5. 1 atom Na **5.** __________

Convert to number of atoms.

6. 540 grams Al **6.** __________

7. 294 grams Au **7.** __________

8. 6.35 grams Cu **8.** __________

9. 2000 grams Mg **9.** __________

10. 1.00 gram Li **10.** __________

10C

Moles of Compounds—Two-Step Problems (continued)

Exercises

Begin each problem by sketching a diagram that outlines the steps in the solution to the problem.

Convert to number of molecules.

1. 72 grams HCl — **1.** ______________
2. 9.0 grams H_2O — **2.** ______________
3. 22 grams CO_2 — **3.** ______________
4. 500 grams NO — **4.** ______________
5. 1.00 gram CCl_4 — **5.** ______________

Convert to mass in grams.

6. 6.02×10^{23} molecules Cl_2 — **6.** ______________
7. 3.01×10^{23} molecules SO_2 — **7.** ______________
8. 1.81×10^{24} molecules CO_2 — **8.** ______________
9. 1000 molecules H_2S — **9.** ______________
10. 1 molecule H_2O — **10.** ______________

10)

Name ____

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Chemistry – Unit 5 Worksheet 1

1. An old (pre-1987) penny is nearly pure copper. If such a penny has a mass of 3.3 g, how many moles of copper atoms would be in one penny?

2. Four nails have a total mass of 4.42 grams. How many moles of iron atoms do they contain?

3. A raindrop has a mass of 0.050 g. How many moles of water does a raindrop contain?

4. What mass of water would you need to have 15.0 moles of H_2O?

5. One box of Morton's Salt contains 737 grams. How many moles of sodium chloride is this?

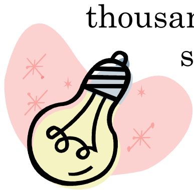
6. A chocolate chip cookie recipe calls for 0.050 moles of baking soda (sodium bicarbonate). How many grams should the chef mass out?

7. Rust is iron(III) oxide. The owner of a l959 Cadillac convertible wants to restore it by removing the rust with oxalic acid, but he needs to know how many moles of rust will be involved in the reaction. How many moles of iron(III) oxide are contained in 2.50 kg of rust?

8. First-century Roman doctors believed that urine whitened teeth and also kept them firmly in place. As gross as that sounds, it must have worked because it was used as an active ingredient in toothpaste and mouthwash well into the 18th century. Would you believe it's still used today? Thankfully, not in its original form! Modern dentists recognized that it was the ammonia that cleaned the teeth, and they still use that. The formula for ammonia is NH_3. How many moles are in 0.75 g of ammonia? How many molecules?

9. Lead (II) chromate, $PbCrO_4$, was used as a pigment in paints. How many moles of lead chromate are in 75.0 g of lead (II) chromate? How many atoms of oxygen are present?

10. The diameter of the tungsten wire in a light bulb filament is very small, less than two thousandths of an inch, or about 1/20 mm. The mass of the filament is so very small – 0.0176 grams – that it would take 1,600 filaments to weigh an ounce! ounce! How many tungsten atoms are in a typical light bulb filament?



11. Two popular antacids tablets are Tums and Maalox. The active ingredient in both of these antacids is calcium carbonate. Tums Regular Strength tablets contain 0.747 g and Maalox tablets contain 0.600 g of calcium carbonate. How many more molecules of calcium carbonate does a Tums provide than a Maalox?

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Chemistry – Unit 5 Worksheet 2
Empirical and Molecular Formulas

Show all your work when solving the following problems. Be sure to include units and label your answer.

1. Find the empirical formula of a compound containing 32.0 g of bromine and 4.9 g of magnesium.

2. What is the empirical formula of a carbon-oxygen compound, given that a 95.2 g sample of the compound contains 40.8 g of carbon and the rest oxygen?

3. A compound was analyzed and found to contain 9.8 g of nitrogen, 0.70 g of hydrogen, and 33.6 g of oxygen. What is the empirical formula of the compound?

4. A compound composed of hydrogen and oxygen is found to contain 0.59 g of hydrogen and 9.40 g of oxygen. The molar mass of this compound is 34.0 g/mol. Find the empirical and molecular formulas.

5. A sample of iron oxide was found to contain 1.116 g of iron and 0.480 g of oxygen. Its molar mass is roughly 5 x as great as that of oxygen gas. Find the empirical formula and the molecular formula of this compound.

6. Find the percentage composition of a compound that contains 17.6 g of iron and 10.3 g of sulfur. The total mass of the compound is 27.9 g.

7. Find the percentage composition of a compound that contains 1.94 g of carbon, 0.48 g of hydrogen, and 2.58 g of sulfur in a 5.00 g sample of the compound.

8. What is the % by mass of oxygen in $Mg(NO_3)_2$?

Molecular and Empirical Formulas

1. Determine the percent composition of the compound that consists of 71.72 g Chlorine, 16.16 g oxygen and 12.12 g carbon.

2. The compound caffeine is often found in soft drinks. A given sample is composed of 9.61 g of carbon, 1.00 g of hydrogen, 5.60 g of nitrogen and 3.2 g of oxygen.
   (a) find its empirical formula

   (b) If its molecular mass is 194.2 g/mole, what is its molecular formula?

3. You find that 3.2 % of a compound is hydrogen, 19.4 % of a compound is carbon and 77.4 % is oxygen.
   (a) Find the empirical formula for the compound

   (b) 2.33 moles of the compound are found to have a mass of 145 g. What is the molecular formula of the compound?

4. A 250 g iron chunk is left out in the air to oxidize . Three years later the iron has changed to a whitish color and the remains of the chunk have a mass of 337 g.
   (a) What is the percent composition of the compound which was formed from the iron bar after the three year period?

   (b) What is the empirical formula of the compound which was formed from the iron bar?

   (c) If the molar mass of the compound is 72.0 g/mole, what is the molecular formula of the compound?

5. A 270.02 g sample of a compound is decomposed in a Hoffman Apparatus (using electricity). After the decomposition, the scientist notes that 70.04 g of nitrogen gas remain in the container and that an unknown amount of oxygen gas has escaped.
   (a) What is the percent composition of the compound which was decomposed?

   (b) What is the empirical formula of the compound before its decomposition?

   (c) If the molar mass of the compound is 108.01 g/mole, what is the molecular formula of the compound?

Chemistry – Unit 5 Review

1. Definitions
   a. mole

   b. molar mass

   c. Avogadro's number

   d. empirical formula

   e. molecular formula

2. Find the molar mass of the following:
   a. KNO_3 ____________________
   b. $(NH_4)_2CO_3$ ____________________
   c. Ag_2CrO_4 ________________
   d. oxygen gas ____________
   e. calcium nitrate ______________
   f. lead(II) nitrate ______________

3. Consider the masses of various hardware below.

Type	Mass (g)	Relative mass
Washer	1.74	
Hex nut	3.16	
Anchor		3.00
Bolt	7.64	

   a. Do the calculations necessary to complete the table.

   b. Explain the connection between these calculations and the atomic masses in the Periodic Table.

4. Convert from g → moles or from moles → g. Show units.

   a. 12.0 g Fe x ______________ = _______________ moles

   b. 25.0 g of chlorine gas x ______________ = _______________ moles

   c. 0.476 g of $(NH_4)_2SO_4$ x ______________ = _______________ moles

   d. 0.15 moles $NaNO_3$ x __________ = _______________ g

   e. 0.0280 moles NO_2 x __________ = _______________ g

   f. 0.64 moles aluminum chloride x ____________ = _______________ g

5. Use Avogadro's number to do the following. Show work, use labels.

   a. How many atoms are there in 0.00150 moles Zn?

   b. If you had 2.50 moles of oxygen gas, what mass of the gas would be in the sample?

   c. A 4.07 g sample of NaI contains how many atoms of Na?

   d. How many atoms of chlorine are there in 16.5 g of iron (III) chloride?

   *e. What is the mass of 100 million atoms of gold? Could you mass this on a balance?

6. Calculate the empirical formula of a compound that contains 4.20 g of nitrogen and 12.0 g of oxygen.

7. When 20.16 g of magnesium oxide reacts with carbon, carbon monoxide forms and 12.16 g of Mg metal remains. What is the empirical formula of magnesium oxide?

8. What is the molecular formula of each compound?

Empirical Formula	Actual Molar Mass of Compound	Molecular Formula
CH	78 g/mole	____________
NO_2	92 g/mole	____________

9. A compound is composed of 7.20 g of carbon, 1.20 g of hydrogen and 9.60 g of oxygen. The molar mass of the compound is 180 g/mole. Determine the empirical and molecular formulas of this compound.

10. What is the % by mass of oxygen in water?

11. A compound of iron and oxygen is found to contain 28 g of Fe and 8.0 g of O. What is the % by mass of each element in the compound?

Chemistry - Nail Lab

Purpose

The purpose is to determine the ratio of copper produced to iron consumed in a single replacement reaction.

Procedure

Day 1

1. Label, then mass a 250 mL beaker.

2. Put between 6.0 and 8.5 g of copper(II) chloride in the beaker. To do this, move the rider on the balance beam up by the value you chose, add copper(II) chloride to beaker until the balance is level again.

3. Add about 50 mL distilled water to the beaker. Stir to dissolve the solid.

4. Mass 2 or 3 nails together to ± 0.01g.

5. Place the nails in the copper chloride solution. Observe the reaction; record your observations. Place the labeled beaker in the place designated by your teacher.

Day 2

6. Remove the nails. Rinse or scrape the precipitate (copper metal) from the nails into your labeled 250 mL beaker. Place the nails in a labeled small beaker. Note the appearance of the nails. Place this beaker in the drying oven.

7. Decant solution from the 250 mL beaker. Rinse the precipitate with about 25 mL of distilled water. Try to lose as little of the solid copper as you can when you decant. After a 2^{nd} rinse with distilled water, rinse the copper with 25 mL of 1 M HCl. Rinse one last time with distilled water. Then place the labeled beaker in the drying oven.

8. Mass the dry nails, then discard the nails.

Day 3

9. Mass the beaker + dry copper. Discard the copper in the place designated by your teacher. Wash your beaker and let dry.

Data:

Mass 250 mL beaker	
Mass 250 mL beaker + copper(II) chloride	
Mass nails before reaction	
Mass nails after reaction	
Mass 250 mL beaker + dry copper	

Calculations:

1. Determine the mass of copper produced and the mass of iron used during the reaction.

2. Calculate the moles of copper and moles of iron involved in the reaction.

3. Determine the ratio $\dfrac{\text{moles of copper.}}{\text{moles of iron}}$

   Express this ratio as an integer. For example, a ratio of 1.33 can be expressed as $\frac{4}{3}$; 0.67 can be expressed as $\frac{2}{3}$, etc.

Conclusion:

1. Why did the reaction stop? Which reactant was used up? How do you know?

2. Describe what was happening to the atoms of iron and copper during the reaction. What is this type of reaction called?

3. What would happen to the ratio of copper to iron if you had placed more nails in the beaker? If you let the reaction go for less time?

4. What is the accepted ratio of copper atoms to iron atoms in this reaction.? Account for differences between your experimental value and the accepted value.
   Write the balanced equation for the reaction.